Chapter 4
Electrolytes
Acid-Base (Neutralization)
Oxidation-Reduction (Redox)
Reactions

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Types of Reactions

Two classifications: one how atoms are rearrangement and the other is chemical reaction

1) Atomic Rearrangement

- **Synthesis** (combination): two substances combine to form one.

  $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

- **Double Displacement**: A reaction in which two elements displaces two elements.

  $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

- **Single displacement**: A reaction where one element displaces one other element.

  $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

- **Decomposition**: A reaction in which a single compound reacts to give two or more substances.

  $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

2) Chemical Classification: Types of Chemical Reactions

Precipitation Reactions: where a solid is formed when two solutions are mixed.

Neutralization Reactions: when an acid and base react to from salt and water.

Oxidation–Reduction Reactions: addition or removal of oxygen and/or transfer of electrons.

Strong Acids

These acids dissociate completely

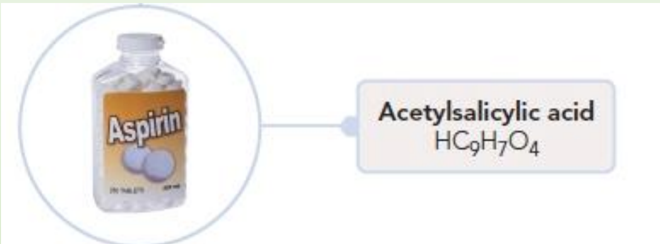
Acid	Ionization Equation
Hydrochloric acid	$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$
Hydrobromic acid	$HBr(aq) \longrightarrow H^+(aq) + Br^-(aq)$
Hydroiodic acid	$HI(aq) \longrightarrow H^+(aq) + I^-(aq)$
Nitric acid	$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$
Chloric acid	$HClO_3(aq) \longrightarrow H^+(aq) + ClO_3^-(aq)$
Perchloric acid	$HClO_4(aq) \longrightarrow H^+(aq) + ClO_4^-(aq)$
Sulfuric acid*	$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$
	$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$

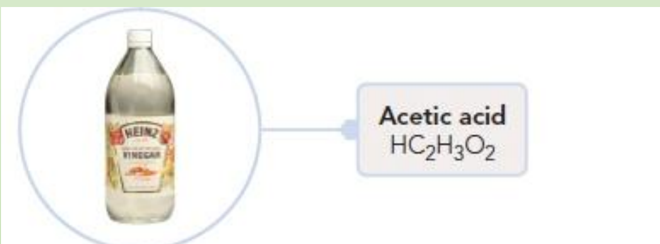
*Note that although each sulfuric acid molecule has two ionizable hydrogen atoms, it only undergoes the first ionization completely, effectively producing one H^+ ion and one HSO_4^- ion per H_2SO_4 molecule. The second ionization happens only to a very small extent.

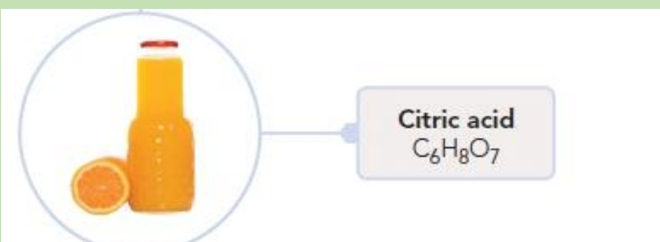
Neutralization Reactions (acid-base)

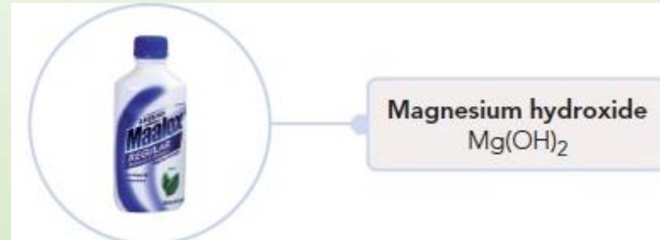
Acids	Bases
Arrhenius Acid A substance that produces hydrogen ions, H^+, when dissolved in water.	**Arrhenius Base** A substance that produces hydroxide ions, OH^-, when dissolved in water.
Brønsted–Lowry Acid The species (molecule or ion) that donates a proton, H^+, to another species in a proton−transfer reaction.	**Brønsted–Lowry Base** The species (molecule or ion) that accepts a proton, H^+, from another species in a proton−transfer reaction.
Sour	Bitter
Corrosive	Caustic, slippery
pH value 1-7	pH value 7-14
Strong acids (inorganic acids) – ionize completely in water, e.g.: HNO_3, H_2SO_4, $HClO_4$, HCl, HBr, HI	Strong bases (inorganic bases) – ionize completely in water; most are hydroxides, e.g.: NaOH, KOH, $Ca(OH)_2$
Weak acids – ionize partially in water, e.g. HF Organic acid: $HC_2H_3O_2$ (CH_3COOH)	Weak bases– ionize partially in water, e.g.: NH_4OH, Na_2CO_3, $NaHCO_3$ organic bases: CH_3NH_2

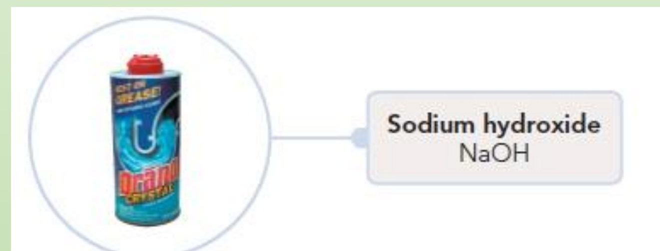
Common Acids and Bases











More on Acids-Bases

Indicators: these are chemicals used to determine if an acid or base is strong or weak by changing colors.

Polyprotic Acid: An acid that results in two or more acidic hydrogens per molecule. E.g. HCl has only one proton to give; but H_2SO_4, sulfuric acid can give 2 protons.

Acid-Base Neutralization Reactions

Neutralization Reaction:

- Almost all acid base reactions are double displacement reactions.
- Most will produce a salt and water as product.
- Carbonates and sulfites give CO_2 and SO_2 gases in product.

Neutralization: Reaction between an acid and a base

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

Molecular equation:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Ionic equation:

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{Na^+}(aq) + OH^-(aq) \rightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$$

Net ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the neutralization of sulfurous acid, H_2SO_3, by potassium hydroxide, KOH

The reaction is a double displacement reaction.

Molecular Equation

(Balance the reaction and include state symbols)

$$H_2SO_3(aq) + 2KOH(aq) \rightarrow 2H_2O(l) + K_2SO_3(aq)$$

Ionic Equation

$$H_2SO_3(aq) + 2\cancel{K^+}(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2\cancel{K^+}(aq) + SO_3^{2-}(aq)$$

Net Ionic Equation

$$H_2SO_3(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + SO_3^{2-}(aq)$$

Neutralization Reactions Producing Gases

Sulfides, carbonates, sulfites react with acid to form a gas.

$Na_2S(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2S(g)$

$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

$Na_2SO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$

The photo below shows baking soda (sodium hydrogen carbonate) reacting with acetic acid in vinegar to give bubbles of carbon dioxide.

+

→

Neutralization Reaction – another one

Molecular Equation

(Balance the reaction and include state symbols)

$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$

Ionic Equation

$CuCO_3(s) + 2H^+(aq) + \cancel{2Cl^-(aq)} \rightarrow Cu^{2+}(aq) + \cancel{2Cl^-(aq)} + H_2O(l) + CO_2(g)$

Net Ionic Equation

$$CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$$

Redox Reactions

Oxidation	Reduction
Addition of oxygen	Removal of oxygen
Removal of hydrogen	Addition of hydrogen
Loss of electrons (LEO)	Gain of electrons (GER)
Metals lose electrons hence undergo oxidation	Non metals gain electrons hence undergo reduction
Reducing agents – something that causes reduction of another element and gets oxidized (loses electrons) itself	Oxidizing agent – an element that causes oxidation of another element and gets reduced (gains electrons) itself

Writing Redox Reactions

Example

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

gaining 2 electrons, reduction

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

loses 2 electrons, oxidation

Half Reactions:

Zinc is losing 2 electrons and oxidized.

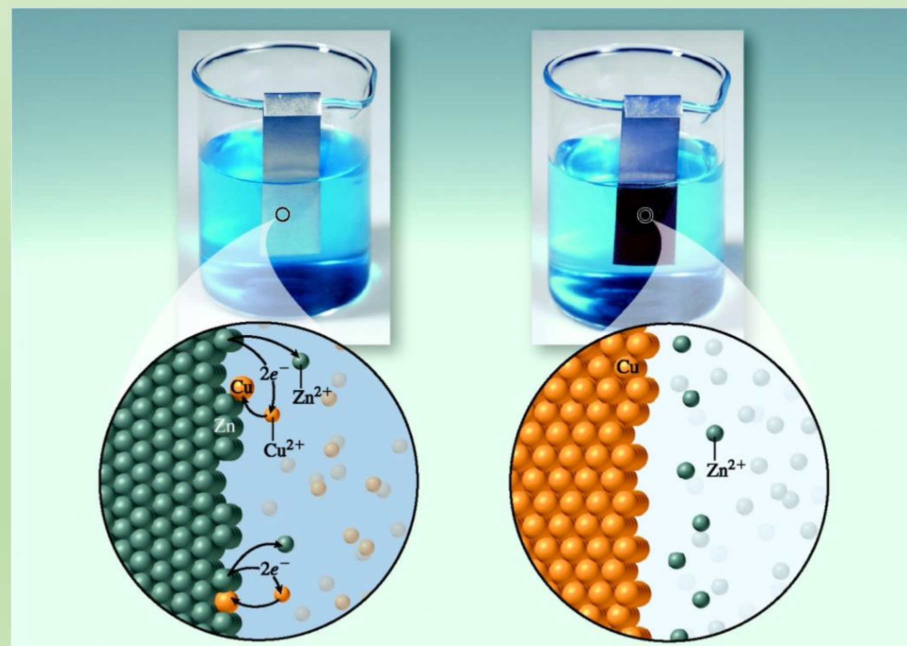
It is the reducing agent

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$$

Copper ions are gaining the 2 electrons.

It is the oxidizing agent

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$$



Rules for Assigning Oxidation Numbers

1. **Elements:** The oxidation number of an element is zero.
2. **Monatomic ions:** The oxidation number of a monatomic ion equals the charge on the ion.
3. **Oxygen:** The oxidation number of oxygen is −2 in most compounds. (An exception is O in H_2O_2 and other peroxides, where the oxidation number is −1.)
4. **Hydrogen:** The oxidation number of hydrogen is +1 in most of its compounds. (The oxidation number of hydrogen is −1 in binary compounds with a metal such as CaH_2.)
5. **Halogens:** The oxidation number of fluorine is −1. Each of the other halogens (Cl, Br, I) has an oxidation number of −1 in binary compounds, except when the other element is another halogen above it in the periodic table or the other element is oxygen.
6. **Compounds and ions:** The sum of the oxidation numbers of a compound is *zero*. The sum of the oxidation numbers of the atoms in a polyatomic ion equals the charge on the ion.

Oxidation Numbers on the Periodic Table

(most common in red)

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H** +1 −1																	2 **He**
3 **Li** +1	4 **Be** +2											5 **B** +3	6 **C** +4 +2 −4	7 **N** +5 +4 +3 +2 +1 −3	8 **O** +2 $-\frac{1}{2}$ −1 −2	9 **F** −1	10 **Ne**
11 **Na** +1	12 **Mg** +2											13 **Al** +3	14 **Si** +4 −4	15 **P** +5 +3 −3	16 **S** +6 +4 +2 −2	17 **Cl** +7 +6 +5 +4 +3 +1 −1	18 **Ar**
19 **K** +1	20 **Ca** +2	21 **Sc** +3	22 **Ti** +4 +3 +2	23 **V** +5 +4 +3 +2	24 **Cr** +6 +5 +4 +3 +2	25 **Mn** +7 +6 +4 +3 +2	26 **Fe** +3 +2	27 **Co** +3 +2	28 **Ni** +2	29 **Cu** +2 +1	30 **Zn** +2	31 **Ga** +3	32 **Ge** +4 −4	33 **As** +5 +3 −3	34 **Se** +6 +4 −2	35 **Br** +5 +3 +1 −1	36 **Kr** +4 +2
37 **Rb** +1	38 **Sr** +2	39 **Y** +	40 **Zr** +4	41 **Nb** +5 +4	42 **Mo** +6 +4 +3	43 **Tc** +7 +6 +4	44 **Ru** +8 +6 +4 +3	45 **Rh** +4 +3 +2	46 **Pd** +4 +2	47 **Ag** +1	48 **Cd** +2	49 **In** +3	50 **Sn** +4 +2	51 **Sb** +5 +3 −3	52 **Te** +6 +4 −2	53 **I** +7 +5 +1 −1	54 **Xe** +6 +4 +2
55 **Cs** +1	56 **Ba** +2	57 **La** +3	72 **Hf** +4	73 **Ta** +5	74 **W** +6 +4	75 **Re** +7 +6 +4	76 **Os** +8 +4	77 **Ir** +4 +3	78 **Pt** +4 +2	79 **Au** +3 +1	80 **Hg** +2 +1	81 **Tl** +3 +1	82 **Pb** +4 +2	83 **Bi** +5 +3	84 **Po** +2	85 **At** −1	86 **Rn**

Activity Series

Loses electrons easily

TABLE 4.6 Activity Series

Increasing ease of oxidation

Element	Oxidation Half-Reaction
Lithium	$Li \longrightarrow Li^{+} + e^{-}$
Potassium	$K \longrightarrow K^{+} + e^{-}$
Barium	$Ba \longrightarrow Ba^{2+} + 2e^{-}$
Calcium	$Ca \longrightarrow Ca^{2+} + 2e^{-}$
Sodium	$Na \longrightarrow Na^{+} + e^{-}$
Magnesium	$Mg \longrightarrow Mg^{2+} + 2e^{-}$
Aluminum	$Al \longrightarrow Al^{3+} + 3e^{-}$
Manganese	$Mn \longrightarrow Mn^{2+} + 2e^{-}$
Zinc	$Zn \longrightarrow Zn^{2+} + 2e^{-}$
Chromium	$Cr \longrightarrow Cr^{3+} + 3e^{-}$
Iron	$Fe \longrightarrow Fe^{2+} + 2e^{-}$
Cadmium	$Cd \longrightarrow Cd^{2+} + 2e^{-}$
Cobalt	$Co \longrightarrow Co^{2+} + 2e^{-}$
Nickel	$Ni \longrightarrow Ni^{2+} + 2e^{-}$
Tin	$Sn \longrightarrow Sn^{2+} + 2e^{-}$
Lead	$Pb \longrightarrow Pb^{2+} + 2e^{-}$
Hydrogen	$H_2 \longrightarrow 2H^{+} + 2e^{-}$
Copper	$Cu \longrightarrow Cu^{2+} + 2e^{-}$
Silver	$Ag \longrightarrow Ag^{+} + e^{-}$
Mercury	$Hg \longrightarrow Hg^{2+} + 2e^{-}$
Platinum	$Pt \longrightarrow Pt^{2+} + 2e^{-}$
Gold	$Au \longrightarrow Au^{3+} + 3e^{-}$

Does not lose electrons easily

Assigning Oxidation States

Assign oxidation numbers for all elements in each species

1) $MgBr_2$: Mg +2, Br −1 x 2= -2; *+2 +(-2)* = total charge of 0

2) ClO_2^- : O −2 x 2 = -4; Cl +3; *(-4) + (+3)* = -1 (charge left over on ion)

3) Assign oxidation number of Mn in K**<u>Mn</u>**O_4

K	**Mn**	**O**	
1(+1)	+ 1(oxidation number of Mn)	+ 4(−2)	= 0
1	+ 1(oxidation number of Mn)	+ (−8)	= 0
	(−7) + (oxidation number of Mn)		= 0

Oxidation number of Mn = +7

4) What is the oxidation number of Cr in dichromate, **<u>Cr</u>**$_2O_7^{2-}$?

Cr	**O**	
2(oxidation number of Cr)	+ 7(−2)	= −2
2(oxidation number of Cr)	+ (−14)	= −2
2(oxidation number of Cr)		= +12

Oxidation number of Cr = +6

Balancing Redox Equations (electronically)

Balance the following reaction.

$$Zn(s) + Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$$

Oxidation Numbers 0 + +2 0

Next, write the unbalanced half–reactions.

$Zn(s) \rightarrow Zn^{2+}(aq)$ (oxidation)

$Ag^{+}(aq) \rightarrow Ag(s)$ (reduction)

Now, balance the charge in each half reaction by adding electrons.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$ (oxidation)

$e^{-} + Ag^{+}(aq) \rightarrow Ag(s)$ (reduction)

Each half–reaction should have the same number of electrons. To do this, multiply each half–reaction by a factor so that when the half–reactions are added, the electrons cancel.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$ (oxidation)

$2e^{-} + 2Ag^{+}(aq) \rightarrow 2Ag(s)$ (reduction)

Lastly, add the two half–reactions together.

$$Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$$

One more Balancing Redox Equation

$$FeI_3(aq) + Mg(s) \rightarrow Fe(s) + MgI_2(aq)$$

The oxidation numbers are given below the reaction.

$$\underset{+3\ \ -1}{FeI_3(aq)} + \underset{0}{Mg(s)} \rightarrow \underset{0}{Fe(s)} + \underset{+2\ \ -1}{MgI_2(s)}$$

Now, write the half–reactions. Since Iodide is a spectator ion it is omitted at this point.

$Mg(s) \rightarrow Mg^{2+}(aq)$ (oxidation)

$Fe^{3+}(aq) \rightarrow Fe(s)$ (reduction)

Balancing the half–reactions:

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (oxidation)

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$ (reduction)

Multiply the oxidation half–reaction by 3 and the reduction half–reaction by 2.

$3Mg(s) \rightarrow 3Mg^{2+}(aq) + 6e^-$ (oxidation)

$2Fe^{3+}(aq) + 6e^- \rightarrow 2Fe(s)$ (reduction)

Add the half–reactions together.

$2Fe^{3+}(aq) + 3Mg(s) \rightarrow 2Fe(s) + 3Mg^{2+}(aq)$

Now, return the spectator ion, I^-.

$2FeI_3(aq) + 3Mg(s) \rightarrow 2Fe(s) + 3MgI_2(aq)$

Types of Redox Reactions

Displacement reactions

A common reaction: active metal replaces (displaces) a metal ion from a solution (use the activity series to predict if reaction will take place)

$$Mg(s) + CuCl_2(aq) \rightarrow Cu(s) + MgCl_2(aq)$$

Decomposition reactions

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Combination Reactions

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Combustion reactions

Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water

$$CH_4(g) + O_2(g) \rightarrow H_2O(l) + CO_2(g)$$

Solved Example

Classify the following reactions as precipitation; acid-base; or redox reaction and any other classification that can describe the reaction.

$2H_2 + O_2 \rightarrow 2H_2O$

Redox (combustion, combination)

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Redox (single displacement)

$H_2O + NH_3 \rightarrow NH_4^+ + OH^-$

Acid-base (double displacement)

$6FeSO_4 + K_2Cr_2O_7 + 7H_2SO_4 \rightarrow Cr_2(SO_4)_3 + 3Fe_2(SO_4)_3 + K_2SO_4 + 7H_2O$

Redox

$2NaCl + Pb(NO_3)_2 \rightarrow PbCl_2 + 3NaNO_3$

Precipitation (double displacement)

Key Words and Concepts

- **Types of Chemical Reactions**
  - Acid–Base Reactions
  - Oxidation–Reduction Reactions
    - Oxidation
    - Reduction
    - Reducing agent
    - Oxidizing agent
    - Half reactions
    - Activity series